



CHEM 150

ACIDS AND BASES

Definitions:

- Bronsted-Lowry
 - Acid – Donate H^+
 - Base – Accept H^+
 - Ex. $HCl + NaOH \rightarrow H_2O + NaCl$
A B
- Lewis
 - Acid – Accept a lone pair
 - Base – Donate a lone pair
 - Ex. $BF_3 + NH_3 \rightarrow BF_3 \cdot NH_3$
A B
- Arrhenius
 - Acid – Cause H^+ to form
 - Base – Cause OH^- to form
 - Ex. $CO_2 + H_2O \rightarrow HCO_3^- + H^+$
A B
- Solvent system
 - Acid – Cause cation to form
 - Base – Cause anion to form
- Lux-Flood
 - Acid – Oxide acceptor
 - Base – Oxide donor
- Note:
 - Neutral metals tend to be neutral or basic (by Arrhenius def)
 - Metal ions tend to act as acids in the sense that they help increase H^+ concentration
 - Hard : more ionic character
 - Soft : more covalent character

Trends in strength:

- The stability of the Conj. Base will determine the strength of the acid
 - Bond dissociation energy is a bigger deciding factor
- Conj. Base
 - Increase size, increase stability
 - Better distribution of charge
 - Increase # of lone pairs, increase stability
 - Better distribution of charge
 - Increase EWG, increase stability
 - Electronegative elements pull e^- toward themselves, making H^+ easier to pull off
- In general, increase # of oxygens in acid, increase in acidity
 - Ex. $H_2SO_4 > H_2SO_3$
- Pauling's Rule: $pK_a \sim 9 - 7n$ where $n = \#$ unprotonated oxygen in neutral
 - Ex. $HClO_4$ (3 unprotonated O)
 - $pK_a \sim 9 - (7)(3) = -12$
- Lower pK_a is the stronger acid

